Stoichiometry − Limiting Reactants – in grams

Name______________________

Again let me reemphasize that there is no substitute for reading the problem carefully. There will be several numbers in each problem, (some of which may not even be necessary) and you must be sure and use the appropriate numbers at the appropriate times. Each number in all of these problems should have three labels associated with it.

1. the units (g, mole, etc)
2. the Identity label – who the substance is (H_2O , carbon dioxide, Al, lead, etc)
3. the Descriptor label: descriptive words to tell you more about the material – (started with, produced, needed, mass theoretically produced, experimental mass, left over, etc)

For you to have success you need to keep track of the labels on every number both at the start of the problem and throughout the problem. If you lose track of who's who, you are likely to use the wrong number at the wrong time.

Problem Solving Plan

As before there is a basic pattern to all stoichiometry problems, with variations depending on what information is given and what questions must be answered. For limiting reactant problems, the problem will give you information about two reactants (as opposed to the one given in the earlier type problems).

A. You must always start with a balanced equation. (If you need more help with that go back to unit G)

B. If it is a limiting reactant problem.... Determine which reactant LIMITS
   - First you *must* change your mass values to moles.
   - The mathematical trick to determine which reactant limits is to divide the moles of each reactant by the coefficient (from the balanced equation) associated with that reactant. The number that comes out the smallest indicates which reactant is the limiting one. The limiting reactant is the one that you must base all your other calculations on because it is the substance that limits how much of everything else can be made or is needed.

C. Answer and questions asked using your Balance Equation Links:
   - Always start with the limiting reactant to answer any questions that are asked.
   - Use whatever stoichiometric LINKS are necessary to convert the moles of the known <u>limiting</u> substance to the moles of one of the desired information.
   - Note that the LINK is set up with the known substance on the bottom (so it will cancel out) and with the desired substance on the top.

D. Of course, the other reactant (if there's only two) will be the excess reactant, and some of it will be left over.

E. Knowing which reactant limits and which is excess, use the limiting reactant to set up a stoichiometric LINK to determine the moles of the excess reactant that is actually needed to do the reaction.

F. Subtract the moles of reactant that you just calculated was needed from the amount of excess reactant started with to determine the moles of excess reactant that is left over.

G. Determining Percent Yield
   - Use the link as instructed in part C to calculate the <u>theoretical</u> amount of the product for which you need a yield.
   - The <u>experimental</u> amount actually produced will be given in the problem. Use it to set up the equation:

$$\frac{\exp erimental}{theoretical} x100 = PercentYield$$

Sample Problems on the next page

Sample Problems – Study these carefully, and model your work after these.

Molar Masses g/mol
$LiOH$ = 23.95
CO_2 = 44.01
Li_2CO_3 = 73.89
H_2O = 18.02

1. Lithium hydroxide is reacted with carbon dioxide to produce solid lithium carbonate and liquid water. If you started with 250.0 g of carbon dioxide and 120.0 g of lithium hydroxide, what mass of water could be made? If you went to the lab and actually produced 35.0 g of H2O, calculate your percent yield. Which reactant is left over, and what mass of it are left over?

   A. Balance the equation: 2 LiOH + CO_2 → Li_2CO_3 + H_2O

   B. Use the "trick" to determine which reactant limits

   - $250gCO_2\left(\frac{1mol}{44g}\right)=\frac{5.68molCO_2}{1}=5$ > $120gLiOH\left(\frac{1mol}{24g}\right)=\frac{5molLiOH}{2}=2.5$ thus LiOH <u>limits the reaction</u>

   C. Determine the mass of product that should be produced (be sure and base you calculations on the limiting reactant).

   - $120gLiOH\left(\frac{1mol}{24g}\right)\left(\frac{1H_2O}{2LiOH}\right)\left(\frac{18g}{1mol}\right)=45.0gH_2O$ <u>should be produced (*theoretical*)</u>

   D. Because the LiOH limits, CO_2 must be the excess reactant.

   - $120gLiOH\left(\frac{1mol}{24g}\right)\left(\frac{1CO_2}{2LiOH}\right)\left(\frac{44g}{1mol}\right)=110gCO_2$ <u>are needed to react with all the LiOH</u>

   E. Calculate how much CO_2 is left over

   - 250 g CO_2 started with – 110 g CO_2 needed = 140 g CO_2 <u>left over</u>

   F. Calculate the percent yield $\frac{35gH_2O}{45gH_2O}x100$ = 77.8% <u>yield of H_2O</u>

Molar Masses in g/mol
C = 12.01
SO_2 = 64.07
CS_2 = 76.15
CO = 28.01

2. Write the reaction that represents the formation of carbon disulfide and carbon monoxide by reacting carbon with sulfur dioxide. If 80.0 g of carbon are reacted with 224.0 g of sulfur dioxide, what mass of carbon monoxide can be produced? What mass of the excess reactant is left over? If 135 g of CO were produced in the lab, what is the percent yield of CO? What if you wanted to use all the left over reactant, how much more of the limiting reactant would you need?

   A. Balance the equation: 5 C + 2 SO_2 → CS_2 + 4 CO

   B. Use the "trick" to determine which reactant limits

   - $80gC\left(\frac{1mol}{12g}\right)=\frac{6.67molC}{5}=1.33$ < $224gSO_2\left(\frac{1mol}{64g}\right)=\frac{3.5molSO_2}{2}=1.75$ thus C <u>limits the reaction</u>

   C. Determine the mass of product that should be produced (be sure and base you calculations on the limiting reactant).

   - $80gC\left(\frac{1mol}{12g}\right)\left(\frac{4CO}{5C}\right)\left(\frac{28g}{1mol}\right)=149gCO$ <u>should be produced (*theoretical*)</u>

   D. Because the C limits, SO_2 must be the excess reactant.

   - $80gC\left(\frac{1mol}{12g}\right)\left(\frac{2SO_2}{5C}\right)\left(\frac{64g}{1mol}\right)=171gSO_2$ <u>are needed to react with all the C</u>

   E. Calculate how much SO_2 is left over

   - 224 g SO_2 started with − 171 g SO_2 needed = 54 g SO_2 <u>left over</u>

   F. Calculate the percent yield $\frac{135gH_2O}{149gCO}x100$ = 90.6% <u>yield of CO</u>

   G. Calculate how much more C would be needed to use up all of the SO_2 that is left over

   - $54gSO_2\left(\frac{1mol}{64g}\right)\left(\frac{5C}{2SO_2}\right)\left(\frac{12g}{1mol}\right)=25gC$ <u>more needed</u> to use up all the excess SO_2

Opener: Stoichiometry & Percent Yield

Worked out on next page

MM(g/mol)
Al = 26.98
$NiSO_4$ = 154.76
Ni = 58.69
$Al_2(SO_4)_3$ = 342.17

Aluminum reacts with nickel(II) sulfate to produce nickel and aluminum sulfate.

1. Write a balanced equation.
2. If 0.0132 mol of aluminum is reacted with 2.65 g of nickel(II) sulfate, what mass of aluminum sulfate could be produced? (Which reactant is the limiting reactant?)
3. If 1.45 g of aluminum sulfate were actually produced in the lab, calculate the percent yield of aluminum sulfate.
4. How many mole of nickel can be produced at the same time?
5. What is the mass of the excess reactant would be left over? (What mole of excess reactant actually reacted?)
6. If you wanted to use up all of the excess reactant, how much more of the limiting reactant would you need to use?

Opener: Stoichiometry & Percent Yield

MM(g/mol)
Al = 26.98
$NiSO_4$ = 154.76
Ni = 58.69
$Al_2(SO_4)_3$ = 342.17

Aluminum reacts with nickel(II) sulfate to produce nickel and aluminum sulfate.

1. Write a balanced equation. $2Al + 3NiSO_4 \rightarrow 3Ni + Al_2(SO_4)_3$
2. If 0.0132 mol of aluminum is reacted with 2.65 g of nickel(II) sulfate, what mass of aluminum sulfate could be produced? (Which reactant is the limiting reactant?) Knowing which reactant Limits (runs out first) is important.

Normalize to determine Limit: $2.65 g\ NiSO_4 \times \frac{1 mol}{154.76 g} = \frac{0.0171 mol\ NiSO_4}{3\ NiSO_4} = 0.0057$

$\frac{0.0132 mol\ Al}{2\ Al} = 0.0066$

Compare — smaller # indicates which reactant Limits

OR calculate amount of product from both reactants

$2.65 g\ NiSO_4 \times \frac{1 mol}{154.76 g} \Rightarrow 0.0171 mol\ NiSO_4 \times \frac{1\ Al_2(SO_4)_3}{3\ NiSO_4} \times \frac{342.17 g}{1 mol} = 1.95 g\ Al_2(SO_4)_3$

$0.0132 mol\ Al \times \frac{1\ Al_2(SO_4)_3}{2\ Al} \times \frac{342.17 g}{1 mol} = 2.26 g\ Al_2(SO_4)_3$

Compare — smaller amount is all that can be made.

3. If 1.45 g of aluminum sulfate were actually produced in the lab, calculate the percent yield of aluminum sulfate. "in the lab" means experimental

$\frac{1.45 g\ Al_2(SO_4)_3\ EXP}{1.95 g\ Al_2(SO_4)_3\ THEOR} \times 100 = 74.2\%$ yield

4. How many mole of nickel can be produced at the same time?

we start this problem with the $NiSO_4$ because it is the limiting reactant (NOT just because it has nickel in it) if Al were the limiting reactant we would need to start with Al.

$0.0171 mol\ NiSO_4 \times \frac{3\ Ni}{3\ NiSO_4} = 0.0171 mol$ Ni can be produced

5. What is the mass of the excess ~~reactant~~ Al would be left over? (What mole of excess ~~reactant~~ Al actually reacted?) we now know the excess is Al

Start with the limit reactant

$0.0171 mol\ NiSO_4 \times \frac{2\ Al}{3\ NiSO_4} =$

0.0132 mol Al Given
− 0.0114 mol Al needed
0.0018 mol Al left over $\times \frac{26.98 g}{1 mol} = 0.0481 g$ Al left

6. If you wanted to use up all of the excess reactant, how much ~~more~~ Mass of the limiting reactant would you need to use?

Got this number from the previous problem

$0.0018 mol\ Al\ left \times \frac{3\ NiSO_4}{2\ Al} \times \frac{26.98 g}{1 mol} = 0.41 g\ NiSO_4$ needed to use up all Al

OR

$0.0132 mol\ Al \times \frac{3\ NiSO_4}{2\ Al} \times \frac{154.76 g}{1 mol} = 3.06 g\ NiSO_4$ need to use up all Al
− 2.65 g $NiSO_4$ given
0.41 g $NiSO_4$ extra needed

Same #!